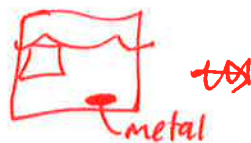


Calorimetry Practice Problem #1



Know

element = metal

$$\text{mass} = 500.0 \text{ g}$$

$$T_i = 153.0^\circ\text{C}$$

$$T_f = 0^\circ\text{C}$$

same!

$$\Delta T = -153.0^\circ\text{C}$$

$$C_p = \frac{26.31 \text{ J}}{\text{mol} \cdot ^\circ\text{C}}$$

$$q = m C_p \Delta T$$

could solve for

$$q_p = n C_p \Delta T$$

want

could solve for

$$q_{\text{metal}} = -q_{\text{ice}}$$

ice-water mixture

$$m = 109.5 \text{ g}$$

$$T_i = 0^\circ\text{C (ice)}$$

$$T_f = 0^\circ\text{C (ice)}$$

$$\Delta T = 0^\circ\text{C, no change}$$

$$\Delta H_{\text{ice}} = \frac{6.02 \text{ kJ}}{\text{mol ice}}$$

want

atomic weight (amu)

??
g/mol
have element I.D. determine

a) Determine q_{ice}

$$\frac{6.02 \text{ kJ}}{\text{mol ice}} \times \frac{109.5 \text{ g ice}}{18.015 \text{ g}} = 36.6 \text{ kJ} = q_{\text{ice}}$$

b) Determine q_{metal}

$$q_{\text{metal}} = -36.6 \text{ kJ}$$

c) Determine moles of metal

$$q_p = n C_p \Delta T$$

$$n = \frac{q_p}{C_p \Delta T}$$

$$= \frac{-36.6 \text{ kJ}}{26.31 \frac{\text{J}}{\text{mol} \cdot ^\circ\text{C}} \times -153.0^\circ\text{C}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 9.09 \text{ mol}$$

$$n = 9.09 \text{ moles metal}$$

d) Determine molar mass

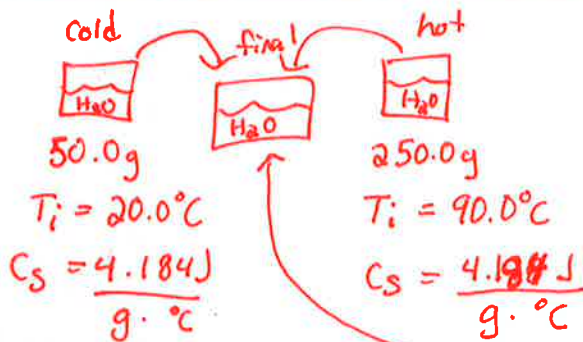
$$\text{molar mass} = \frac{500.0 \text{ g}}{9.09 \text{ mol}} = 55.0 \text{ g/mol}$$

e) Atomic weight & identity

55.0 amu $\rightarrow$ Mn manganese
a metal so makes sense!

Calorimetry Practice Problem #2

Know



Want

T_f sol'n

$$q_{\text{hot}} = -q_{\text{cold}}$$

releases heat

gains heat

equal in magnitude/opposite in sign

$$q = m C_s \Delta T$$

↑ solve for

$$\Delta T = T_f - T_i$$

* Note: There are two ways to solve this. I will show both. Pick the one that makes the most sense to you.

1st way

$$q_{\text{hot}} = -q_{\text{cold}} \text{ so}$$

$$m_{\text{hot}} C_{s,\text{hot}} \Delta T_{\text{hot}} = -m_{\text{cold}} C_{s,\text{cold}} \Delta T_{\text{cold}}$$

these are the same so cancel!

$$(250.0\text{g})(T_f - 90.0^\circ\text{C}) = -(50.0\text{g})(T_f - 20.0^\circ\text{C})$$

$$250.0\text{g} T_f - 22,500^\circ\text{C} = -50.0\text{g} T_f + 1.00 \times 10^3^\circ\text{C}$$

$$+50.0\text{g} T_f \quad +50.0\text{g} T_f$$

$$300.0\text{g} T_f - 22,500^\circ\text{C} = 1.00 \times 10^3^\circ\text{C}$$

$$+ 22,500^\circ\text{C} \quad + 22,500^\circ\text{C}$$

$$300.0\text{g} T_f = 23,500^\circ\text{C}$$

$$\frac{300.0\text{g}}{300.0\text{g}} \quad \frac{23,500^\circ\text{C}}{300.0\text{g}}$$

$$T_f = 78.3^\circ\text{C}$$

* makes sense b/c in between hot & cold temperatures

2nd way

1) Solve for q_{hot}

$$q_{\text{hot}} = m_{\text{hot}} C_{s,\text{hot}} \Delta T_{\text{hot}}$$

$$q_{\text{hot}} = (250.0\text{g}) \left(\frac{4.184\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (T_f - 90.0^\circ\text{C})$$

$$q_{\text{hot}} = \left(\frac{1046\text{J}}{^\circ\text{C}} \right) (T_f - 90.0^\circ\text{C})$$

$$q_{\text{hot}} = \frac{1046\text{J}}{^\circ\text{C}} T_f - 94140^\circ\text{C}$$

$$q_{\text{hot}} = -q_{\text{cold}} \text{ \& } q_{\text{cold}} = m_{\text{cold}} C_{s,\text{cold}} \Delta T_{\text{cold}}$$

2) Substitute in q_{hot} & solve for T_f

$$\frac{1046\text{J}}{^\circ\text{C}} T_f - 94140^\circ\text{C} = -(50.0\text{g}) \left(\frac{4.184\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (T_f - 20.0^\circ\text{C})$$

$$= \left(\frac{-209.2\text{J}}{^\circ\text{C}} \right) (T_f - 20.0^\circ\text{C})$$

$$\frac{1046\text{J}}{^\circ\text{C}} T_f - 94140^\circ\text{C} = \frac{-209.2\text{J}}{^\circ\text{C}} T_f + 4184^\circ\text{C}$$

$$+ \frac{209.2\text{J}}{^\circ\text{C}} T_f \quad + \frac{209.2\text{J}}{^\circ\text{C}} T_f$$

$$\frac{1255.2\text{J}}{^\circ\text{C}} T_f - 94140^\circ\text{C} = 4184^\circ\text{C}$$

(cont)

$$\frac{1255.2 \text{ J}}{^{\circ}\text{C}} T_f = \frac{98324 \text{ J}}{1255.2 \text{ J}}^{\circ}\text{C}$$

$$T_f = 78.3^{\circ}\text{C}$$

* makes sense b/c
in b/w hot & cold
temperatures *

Calorimetry Practice Problem #3

Know

Copper ← only one thing

$$C_s = 0.385 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$$

$$\text{mass} = 16.8 \text{ g}$$

$$T_i = -12.2^\circ\text{C}$$

$$T_f = 25.6^\circ\text{C}$$

$$q = m C_s \Delta T$$

want

heat $\rightarrow q$

units $\rightarrow \text{J or kJ}$

so no system & surroundings

$$\Delta T = T_f - T_i = (25.6^\circ\text{C}) - (-12.2^\circ\text{C}) = 37.8^\circ\text{C} = \Delta T$$

$$q = (16.8 \text{ g}) \left(\frac{0.385 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \right) (37.8^\circ\text{C}) = 2444.9 \text{ J}$$

all have 3 sig. figs. so

2440 J - or -
2.44 x 10³ J - or -
2.44 kJ
all are correct!

Practice Problem #4

Know



$$\text{mass} = 108 \text{ g}$$

$$T_i = 22.5^\circ\text{C}$$



$$T_f = 47.9^\circ\text{C}$$



$$\text{mass} = 65.1 \text{ g}$$

$$T_i = ? \text{ want}$$

← b/c T_f is hotter than T_i of 1st soln
 $q_{\text{cold}} = -q_{\text{hot}}$

$$m_{\text{cold}} C_{s,\text{cold}} \Delta T_{\text{cold}} = - m_{\text{hot}} C_{s,\text{hot}} \Delta T_{\text{hot}}$$

have have have have have

$$m_{\text{cold}} \Delta T_{\text{cold}} = - m_{\text{hot}} (T_{f,\text{hot}} - T_{i,\text{hot}})$$

$$(108 \text{ g}) (47.9^\circ\text{C} - 22.5^\circ\text{C}) = - (65.1 \text{ g}) (47.9^\circ\text{C} - T_i)$$

$$(108 \text{ g}) (25.4^\circ\text{C}) = -3118.3 \text{ g} \cdot ^\circ\text{C} + 65.1 \text{ g} T_i$$

$$\frac{5861.5 \text{ g} \cdot ^\circ\text{C}}{65.1 \text{ g}} = \frac{65.1 \text{ g} T_i}{65.1 \text{ g}}$$

$$90.0^\circ\text{C} = T_i$$

← makes sense b/c should be hotter than final temperature & below the boiling point of water (100°C) or they would have said 'steam' sample.

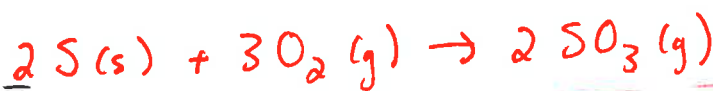
Calorimetry Practice Problem #5

Know

$$\Delta H^\circ_{\text{rxn}} = -790 \text{ kJ}$$

want

ΔH° for this reaction w/
this amount in kJ



0.95 g

1) Convert to moles

$$0.95 \text{ g S} \times \frac{\text{mol S}}{32.07 \text{ g S}} = 0.0296 \text{ mol S}$$

2) Use enthalpy & molar ratios from balanced chemical eq

$$\Delta H^\circ_{\text{rxn}} = \frac{-790 \text{ kJ}}{2 \text{ mol S}} \quad \& \quad \frac{-790 \text{ kJ}}{3 \text{ mol O}_2} \quad \& \quad \frac{-790 \text{ kJ}}{2 \text{ mol SO}_3}$$

$$\text{so } 0.0296 \text{ mol S} \times \frac{-790 \text{ kJ}}{2 \text{ mol S}} = -11.7 \text{ kJ} = \boxed{-12 \text{ kJ}}$$